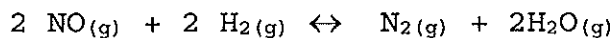


## The Equilibrium Constant: Practice

1. Consider the following reaction:



A mixture at 300K initially contains 0.10mol of NO, 0.050mol of H<sub>2</sub>, and 0.10mol of H<sub>2</sub>O in a 1.0L vessel. At equilibrium, the concentration of NO is 0.062M. What are the equilibrium concentrations of H<sub>2</sub>, N<sub>2</sub>, and H<sub>2</sub>O and the value of the equilibrium constant?

$$K = \frac{[\text{N}_2][\text{H}_2\text{O}]^2}{[\text{NO}]^2[\text{H}_2]^2}$$

$$K = \frac{(0.019)(0.138)^2}{(0.062)^2(0.012)^2}$$

$$K = 650$$

	[NO]	[H <sub>2</sub> ]	[N <sub>2</sub> ]	[H <sub>2</sub> O]
I	0.10	0.050	0	0.10
C	-0.038	-0.038	+0.019	+0.038
E	0.062	0.012	0.019	0.138

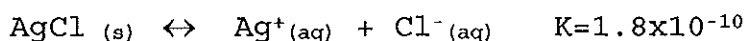
$$[\text{NO}] = 0.062 \text{ M}$$

$$[\text{H}_2] = 0.012 \text{ M}$$

$$[\text{N}_2] = 0.019 \text{ M}$$

$$[\text{H}_2\text{O}] = 0.14 \text{ M}$$

2. Consider the dissolution of silver chloride, AgCl, in water:



Based on the value of K, would you consider AgCl a soluble or insoluble salt? Write the equilibrium expression and determine the [Ag<sup>+</sup>] at equilibrium.

$K \ll 1$  Reactants are favored  $\therefore$  Insoluble salt

$$K = [\text{Ag}^+][\text{Cl}^-] = 1.8 \times 10^{-10}$$

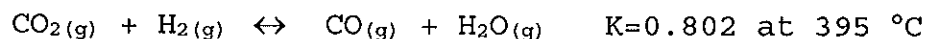
$$(x)(x) = 1.8 \times 10^{-10}$$

$$x = \sqrt{1.8 \times 10^{-10}}$$

$$x = \underline{1.3 \times 10^{-5} \text{ M}} = [\text{Ag}^+]$$

	[AgCl]	[Ag <sup>+</sup> ]	[Cl <sup>-</sup> ]
I	/	0	0
C	/	+x	+x
E	/	x	x

3. Consider the following reaction:



If a reaction mixture contains initially 1.50mol of  $\text{CO}_2$  and 1.50mol of  $\text{H}_2$  are placed in a 0.750L container, what are the masses of  $\text{CO}_2$ ,  $\text{H}_2$ ,  $\text{CO}$ , and  $\text{H}_2\text{O}$  in the equilibrium mixture?

$$K = \frac{[\text{CO}][\text{H}_2\text{O}]}{[\text{CO}_2][\text{H}_2]} = 0.802$$

$$\frac{(x)(x)}{(2.00-x)(2.00-x)} = 0.802$$

$$\frac{x^2}{(2.00-x)^2} = 0.802$$

$$\frac{x}{2.00-x} = \sqrt{0.802}$$

$$x = \sqrt{0.802} (2.00-x)$$

$$x = (\sqrt{0.802})(2.00) - \sqrt{0.802} x$$

$$x = 0.945 \text{ M}$$

	$[\text{CO}_2]$	$[\text{H}_2]$	$[\text{CO}]$	$[\text{H}_2\text{O}]$
I	2.00	2.00	0	0
C	-x	-x	+x	+x
E	2.00-x	2.00-x	x	x

Initial concentrations:

$$[\text{CO}_2] = [\text{H}_2] = \frac{1.50 \text{ mol}}{0.750 \text{ L}} = 2.00 \text{ M}$$

Equil. concentrations:

$$[\text{CO}_2] = [\text{H}_2] = 2.00 \text{ M} - 0.945 \text{ M} = 1.06 \text{ M}$$

$$[\text{CO}] = [\text{H}_2\text{O}] = 0.945 \text{ M}$$

$$\text{mol CO}_2 = \text{mol H}_2 = (1.06 \text{ M})(0.750 \text{ L}) = 0.795 \text{ mol}$$

$$\text{mol CO} = \text{mol H}_2\text{O} = (0.945 \text{ M})(0.750 \text{ L}) = 0.709 \text{ mol}$$

$$\text{mass CO}_2 = 0.795 \text{ mol} \left( \frac{44.01 \text{ g}}{1 \text{ mol}} \right) = 35.0 \text{ g CO}_2$$

$$\text{mass H}_2 = 0.795 \text{ mol} \left( \frac{2.02 \text{ g}}{1 \text{ mol}} \right) = 1.61 \text{ g H}_2$$

$$\text{mass CO} = 0.709 \text{ mol} \left( \frac{28.01 \text{ g}}{1 \text{ mol}} \right) = 19.9 \text{ g CO}$$

$$\text{mass H}_2\text{O} = 0.709 \text{ mol} \left( \frac{18.02 \text{ g}}{1 \text{ mol}} \right) = 12.8 \text{ g H}_2\text{O}$$